Name ____ AP Chemistry

Oxidation-Reduction Reactions

Review of Oxidation-Reduction Reactions

An oxidation-reduction reaction is a reaction in which one or more electrons are transferred. To identify the transfer between species, the oxidation states have to be compared. An oxidation state is an imaginary charge assigned by certain rules that provides a manner to keep track of electrons in a redox reaction. The following list gives the rules for assigning oxidation states:

- Any species in its elemental form (i.e. Na(s), O_2 (g), O_3 (g), Hg (l)) is zero
- A monoatomic ion (i.e. Na^+, Cl^-) is the same as its charge
- Fluorine is -1 in its compounds
- Oxgen is usually -2 in its compounds
  - In peroxides (O_2^{2-}), oxygen is -1
- Hydrogen is +1 in its covalent compounds

Example #1

Assign oxidation states to all of the atoms in the following:

(a) CO_2
(b) SF_4
(c) NO_3^-

Therefore, oxidation is an increase in oxidation state or a loss of electrons while reduction is a decrease in oxidation state or a gain of electrons. Additionally, a species that is reduced and causes another species to be oxidized is called an oxidizing agent. A species that is oxidized and causes another species to be reduced is called a reducing agent.

Example #2

Metallurgy, the process of producing a metal from its ore, always involves oxidation-reduction reactions. In the metallurgy of galena (PbS), the principal lead-containing ore, the first step is the conversion of lead sulfide to its oxide:

$$2\ PbS\ (s) + 3\ O_2\ (g) \rightarrow 2\ PbO\ (s) + 2\ SO_2\ (g)$$

Identify the atoms that are oxidized and reduced, and specify the oxidizing and reducing agents.

Balancing Oxidation-Reduction Reactions

Oxidation-reduction reactions often are complicated reactions so balancing them in a similar manner as other types of reaction is very difficult. A specific method, known as a half-reaction method, has been developed to balance these types of reaction. Additionally, the acidity or basicity of the environment affects an oxidation-reduction reaction so it has to be considered when balancing a redox reaction. First we are going to discuss how to balance these reactions in an acidic environment. Below are the steps for balancing a redox equation in an acidic solution:

1) Write separate half-reactions
2) Balance all atoms except H and O.
3) Balance oxygen using water.
4) Balance hydrogen using H^+.
5) Balance charge using electrons.
6) If necessary, equalize electron transfer by multiplying each half-reaction by numbers to allow for the same number of electrons in each half-reaction.
7) Combine the two reactions and then double check each atoms to ensure the same number is on both sides of the reaction.

<u>Example #3</u>

Potassium dichromate ($K_2Cr_2O_7$) is a bright orange compound that can be reduced to a blue-violet solution of Cr^{3+} ions. Under certain conditions, $K_2Cr_2O_7$ reacts with ethyl alcohol (C_2H_5OH):

$$H^+ \text{ (aq)} + Cr_2O_7^{2-} \text{ (aq)} + C_2H_5OH \text{ (l)} \rightarrow Cr^{3+} \text{ (aq)} + CO_2 \text{ (g)} + H_2O \text{ (l)}$$

Balance this equation using the half-reaction method.

Next we will look at redox reactions in a basic environment. The following steps are for balancing these reactions in basic solutions:

1) Write separate half-reactions
2) Balance all atoms except H and O.
3) Balance oxygen using water.
4) Balance hydrogen using H^+.
5) Balance charge using electrons.
6) If necessary, equalize electron transfer by multiplying each half-reaction by numbers to allow for the same number of electrons in each half-reaction.
7) Add OH^- to balance the H^+ ions.
8) Combine the two reactions and then double check each atoms to ensure the same number is on both sides of the reaction.

Example #4

Silver is sometimes found in nature as large nuggets; more often it is found mixed with other metals and their ores. An aqueous solution containing cyanide ion is often used to extract the silver using the following reaction that occurs in basic solution:

$$Ag\ (s) + CN^-\ (aq) + O_2\ (g) \rightarrow Ag(CN)_2^-\ (aq)$$

Balance this equation using the half-reaction method.

Practice Problems

1) Assign the oxidation states to each element in
   (a) HIO_3
   (b) $NaMnO_4$
   (c) SnO_2
   (d) NOF
   (e) N_2H_4

2) Write each unbalanced half-reaction, label each as oxidation or reduction and then identify the reducing agent and oxidizing agent.
   (a) Ag (s) + NO_3^- (aq) → Ag^+ (aq) + NO (g)
   (b) CO_2 (g) + H_2O → C_2H_4 (g) + O_2 (g)

3) Classify each half-reaction as oxidation or reduction and then balance.
   (a) Mn^{2+} (aq) → MnO_4^- (aq) (acidic)
   (b) CrO_4^{2-} (aq) → Cr^{3+} (aq) (basic)
   (c) PbO_2 (s) → Pb^{2+} (aq) (basic)

4) Write in the balanced equation for each of the following reaction according to the appropriate environment:
   (a) Ni^{2+} (aq) + IO_4^- (aq) → Ni^{3+} (aq) + I^- (aq) (acidic)
   (b) Ca (s) + $Cr_2O_7^{2-}$ (aq) → Ca^{2+} (aq) + Cr^{3+} (aq) (acidic)
   (c) $Fe(OH)_3$ (s) + Cr^{3+} (aq) → $Fe(OH)_2$ (s) + $Cr_2O_7^{2-}$ (aq) (basic)
   (d) MnO_4^- (aq) + BrO_3^- (aq) → MnO_2 (s) + BrO_4^- (aq) (basic)